# Chemistry I: Atomic Structure II: Electron configuration Rules

From quantum mechanics we have a probable location of the electrons in an atom. The four quantum numbers describe an electron. The **Pauli Exclusion Principle** says that no **two electrons** can have the same exact quantum numbers. This means that they **cannot be in the exact same place at the exact same time.**

Electrons are in regions called atomic orbitals. These 3-D regions of space are probable locations of electrons. Any one atomic orbital can only hold 2 electrons. We also know that the orbitals have shapes and are oriented around the X, Y, and Z-axis.

Atomic orbitals are subdivided into suborbitals. This is how it looks:

| Orbital | Suborbitals | # electrons | Labels | shape |
|---|---|---|---|---|
| s | 1 | 2 | | Sphere |
| p | 3 | 6 | $p_x$ $p_y$ $p_z$ | Dumbbells |
| d | 5 | 10 | Too complex | Multiple lobes |
| f | 7 | 14 | Don' t go there | Multiple lobes |

**Aufbau principle**
When **electrons** are **added** to the outside of a nucleus they will **go into the orbital with the lowest energy.** You cannot add outside in.

**Hund's rule**
**Orbitals of equal energy are each occupied by one electron before any one orbital is occupied by a second electron** and all electrons in a singly occupied orbital must have the same spin (must be going the same way as the others). Explanation: when putting electrons in the p-orbitals there must be one electron in $p_x$, $p_y$ and $p_z$ before any of the others can have get a second electron.

**Following the flow chart:**

When writing electron configuration you will need to follow the order of fill chart. This is a systematic way of arranging the electrons. You will need to know how many electrons are in each of the orbitals. The configuration you write may not match what is in Chemistry textbooks because electrons in more complex atoms will move to create more stable configurations based on the lowest possible energy.

**Remember that the most stable electron configuration is that of the noble gases which all have $nS^2nP^6$ configuration.**

## Writing the electron configuration for ions.

Remember **ions are atoms of an element with a charge.**

The charge is due to the number of electrons compared to the number of protons.

So you will need to figure the exact number of electrons.

Examples:

The **neutral sodium** atom has **10 protons and 11 electrons**. The electron configuration for this atom would be $\mathbf{1s^2 2s^2 2p^6 3s^1}$
The **sodium ion ($Na^{+1}$)** has **10 $p^+$ and 10$e^-$** and configuration for this ion is $\mathbf{1s^2 2s^2 2p^6}$

**A neutral atom of fluorine** has **9 $p^+$ and 9 $e^-$**. The configuration for this atom is $\mathbf{1s^2 2s^2 2p^5}$**.** The **fluoride ion ($F^{-1}$)** has **9 $p^+$ and 10 $e^-$** and the configuration is $\mathbf{1s^2 2s^2 2p^6}$

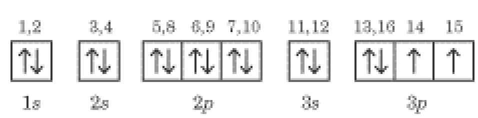
Notice that **both ions** now have the **same configuration**; $\mathbf{1s^2 2s^2 2p^6}$.This is the same configuration for the Noble gas, **neon** which is an extremely stable element and will not react with anything.

**Anions** are atoms or groups of atoms that have *gained* electrons. Having more negatively charged electrons than positively charged protons, they are negatively charged.

**Cations** are atoms or groups of atoms that have *lost* electrons. Having more positively charged protons than negatively charged electrons, they are positively charged.

### Orbital Notation

**Orbital notation** is basically just another way of expressing the electron configuration of an atom. It is very useful in determining quantum numbers as well as electron pairing. The orbital notation for sulfur would be represented as follows:




Notice that electrons 5, 6, and 7 went into their own orbitals before electrons 8, 9, and 10 entered, forcing pairings in the 2*p* sublevel; the same thing happens in the 3*p* level.

Orbital notation is used to diagram the electrons in their paths within an orbital

Examples of ground state electron configurations in the orbital box notation that shows electron spins.

| atom | orbital box diagram | | | | |
|---|---|---|---|---|---|
| B | [↓↑] 1s | [↓↑] 2s | [↑][ ][ ] 2p | | |
| C | [↓↑] 1s | [↓↑] 2s | [↑][↑][ ] 2p | | |
| N | [↓↑] 1s | [↓↑] 2s | [↑][↑][↑] 2p | | |
| O | [↓↑] 1s | [↓↑] 2s | [↓↑][↑][↑] 2p | | |
| F | [↓↑] 1s | [↓↑] 2s | [↓↑][↓↑][↑] 2p | | |
| Cl | [↓↑] 1s | [↓↑] 2s | [↓↑][↓↑][↓↑] 2p | [↓↑] 3s | [↓↑][↓↑][↑] 3p |
| Mn | [↓↑] 1s | [↓↑] 2s | [↓↑][↓↑][↓↑] 2p | [↓↑] 3s | [↓↑][↓↑][↓↑] 3p |
| | | … | [↑][↑][↑][↑][↑] 3d | [↓↑] 4s | |

- drawing orbital box diagrams
  1. write the electron configuration in aufbau order
  2. draw a box for each orbital.
     - Remember that s, p, d, and f orbitals contain 1, 3, 5, and 7 orbitals, respectively.
     - Remember that an suborbital can hold 0, 1, or 2 electrons only, and if there are two electrons in the orbital, they must have opposite (paired) spins (Pauli principle)
  3. within an orbital (depicted as a group of boxes), spread the electrons out and line up their spins as much as possible (Hund's rule)

## Lewis electron dot diagrams: Atoms

1) The symbol of an element is used to represent its nucleus and everything else, save its outer electrons.

2) The first step in drawing a Lewis Dot Diagram is determining the look of the element's outer electrons. For this task, one must figure out the element's electron configuration.

3) From an element's electron configuration, it is possible to identify the electrons in the outer energy level.

4) Every symbol has four sides (top, bottom, left, right). Every side represents an orbital and every orbital can hold up to two electrons. Dots are drawn on appropriate sides representing the electrons in the orbitals.

5) The greatest number of electrons that can be depicted in a Lewis Dot Diagram is eight. Only two of these electrons, those in the "S" orbital, are paired together. This means that they will always be drawn on the same side, although which side that is does not matter.

6) All other electrons are to be distributed as equally as possible around the symbol. This means that there can never be two electrons, other than "S" electrons, on one side while another side is empty.